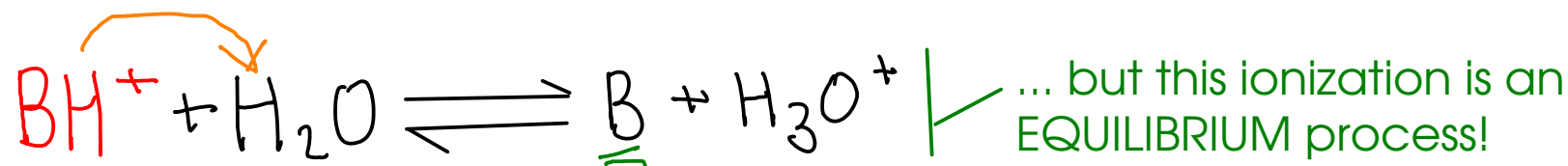
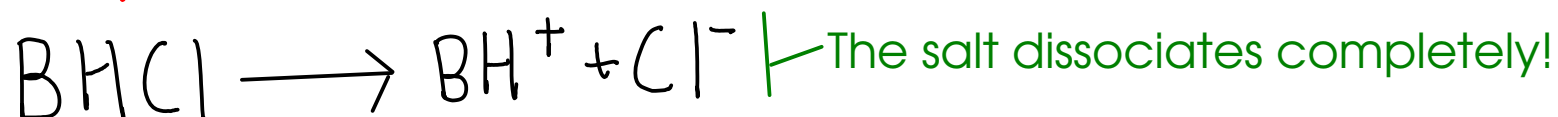


SALT OF A WEAK BASE

ex: NH_4Cl 

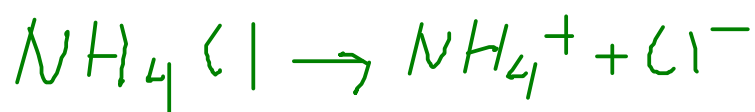
$$K_a = \frac{[\text{B}][\text{H}_3\text{O}^+]}{[\text{BH}^+]} \quad \left| \text{Acid ionization constant for } \text{BH}^+ \right.$$

$$K_w = (K_{a,\text{BH}^+})(K_{b,\text{B}})$$

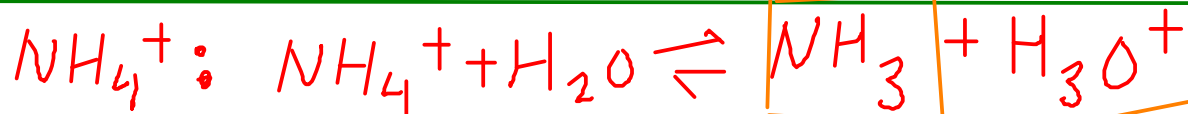
1.0×10^{-14}

Find the pH for salt solutions just like you would find pH for any other weak acid or weak base solutions. Only trick is to find out whether the salt is actually acidic or basic!

0.100 M NH_4Cl ... Find the pH of the solution



Is the salt acidic, basic, or neutral? Look at the IONS the salt releases when it dissolves!



Since ammonia is a WEAK base, that means it will exist in water as molecules ... meaning the reaction we wrote can occur and ammonium ion can act as an acid.



HCl is a strong acid. Since it completely ionizes, that means chloride ion isn't able to keep a proton attached, and isn't a base.

So the equilibrium we'll need to look at is the one with the acidic ammonium ion.



$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} = ?$$

$K_{a, \text{NH}_4^+} = ?$ We'll need to find K_a for ammonium by looking at K_b for its conjugate, ammonia.

$$K_{b, \text{NH}_3} = 1.8 \times 10^{-5} \text{ (p}K_b = 14)$$

$$K_a \times K_b = 1.0 \times 10^{-14}$$

$$K_{a, \text{NH}_4^+} = 5.56 \times 10^{-10}$$



Set up and solve this equilibrium like any other weak acid problem!

Species	[Initial]	Δ	[Equilibrium]
NH_3	0	+x	x
H_3O^+	0	+x	x
NH_4^+	0.100	-x	0.100 - x

Let "x" equal the change in ammonia concentration...

$$\frac{(x)(x)}{(0.100 - x)} = 5.56 \times 10^{-10}$$

$$\frac{x^2}{0.100 - x} = 5.56 \times 10^{-10}$$

$x \ll 0.100,$
 \downarrow so $0.100 - x \approx 0.100$

$$\frac{x^2}{0.100} = 5.56 \times 10^{-10}$$

$$x = 7.45 \times 10^{-6} \text{ M} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log_{10}(7.45 \times 10^{-6})$$

$$\boxed{\text{pH} = 5.13}$$

For comparison,

pH = 1.00 for a strong acid like HCl

pH = 2.17 for 0.10 M nitrous acid

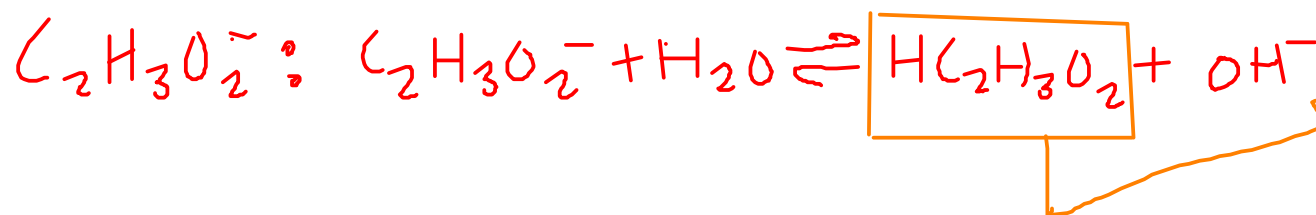
pH = 7.00 for distilled water

0.100 M $\text{NaC}_2\text{H}_3\text{O}_2$, Find pH

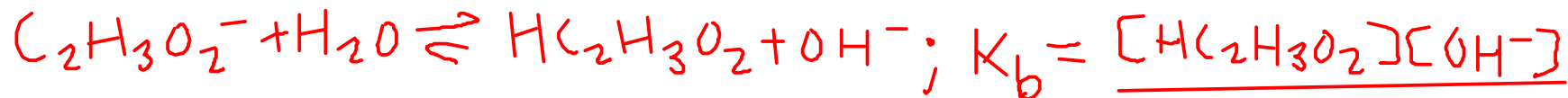


$\text{NaC}_2\text{H}_3\text{O}_2 \rightarrow \text{Na}^+ + \text{C}_2\text{H}_3\text{O}_2^-$ Look at the IONS to find whether the sodium acetate is acidic, basic, or neutral.

Na^+ : Cannot be a B-L acid, since it has no "H" to donate. Since it has a positive charge, it's not likely to be a base, either!



Acetic acid is a WEAK acid, so it's water-stable, meaning that acetate ion can hold on to a proton and be BASIC.



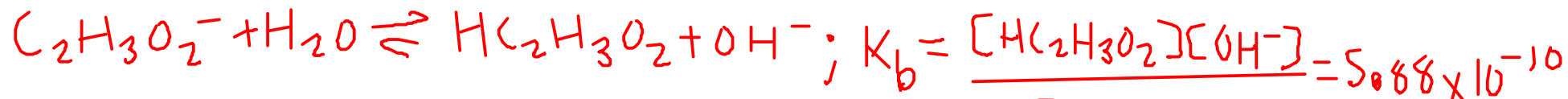
We need a value for K_b for the acetate ion. But it's not in our chart. However, we DO find the value for K_a of acetic acid on page A-13. Convert ...

$$K_a \times K_b = 1.0 \times 10^{-14}$$

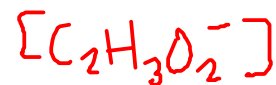
$$(1.7 \times 10^{-5}) K_b = 1.0 \times 10^{-14}$$

$$K_b = 5.88 \times 10^{-10}$$

$$K_a, \text{HC}_2\text{H}_3\text{O}_2 = 1.7 \times 10^{-5}$$



Now solve like any other weak base equilibrium problem!



Species	[Initial]	Δ	[Equilibrium]
$\text{HC}_2\text{H}_3\text{O}_2$	0	+X	X
OH^-	0	+X	X
$\text{C}_2\text{H}_3\text{O}_2^-$	0.100	-X	0.100 - X

Let "x" equal the change in acetic acid concentration

$$\frac{(x)(x)}{(0.100-x)} = 5.58 \times 10^{-10}$$

$$\frac{x^2}{0.100-x} = 5.58 \times 10^{-10}$$

$$\begin{array}{l} x \ll 0.100, \text{ so} \\ \downarrow \\ 0.100 - x \approx 0.10 \end{array}$$

$$\frac{x^2}{0.100} = 5.58 \times 10^{-10}$$

$$x = 7.67 \times 10^{-6} \text{ M} = [\text{OH}^-]$$

$$\text{pOH} = -\log_{10}(7.67 \times 10^{-6})$$

$$\text{pOH} = 5.12$$

$$\text{pH} + \text{pOH} = 14.00$$

$$\text{pH} = 14.00 - 5.12$$

$$\boxed{\text{pH} = 8.88}$$

For comparison:

0.100 M sodium acetate, pH = 8.88

0.100 M ammonia, pH = 11.13

0.100 M NaOH (strong base), pH = 13.00

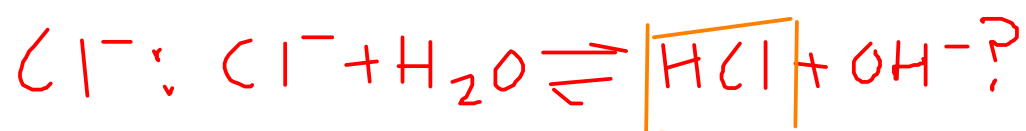
The acetate ion is basic, but it's a very weak base!

0.100 M NaCl, Find pH



To find out if NaCl is acidic, basic, or neutral ... look at the ions.

Na^+ : Can't be a B-L acid, as it has no "H" to donate. The sodium ion's positive charge makes it unlikely to be a proton acceptor. We expect sodium ion to be neutral.



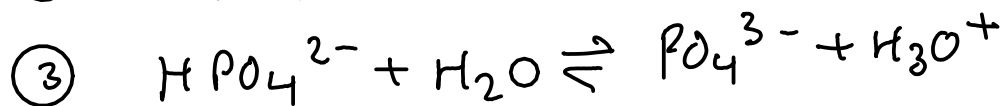
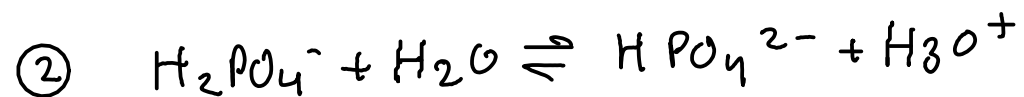
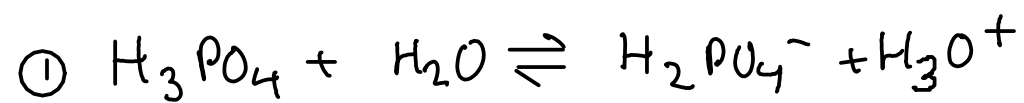
Hydrochloric acid is STRONG (completely ionized), so we don't expect chloride ion to be able to hold on to a proton and function as a base in water. Chloride ion should also be neutral.

So if both sodium ion and chloride ion are neutral, what sets the pH of a 0.100 M NaCl solution?

The water equilibrium itself will control pH here, and the pH will be 7.00

Find pH of 0.10 M H_3PO_4

... what's special about phosphoric acid?



Phosphoric acid has THREE acidic protons!

$$K_{a1} = 6.9 \times 10^{-3}$$

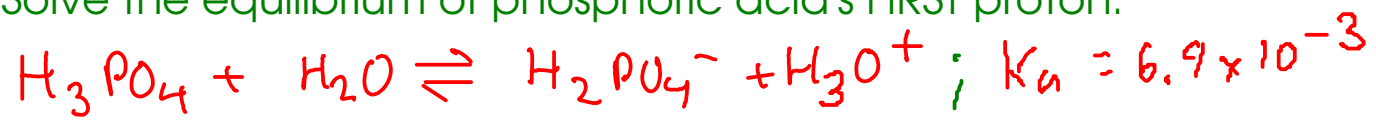
$$K_{a2} = 6.2 \times 10^{-8}$$

$$K_{a3} = 4.8 \times 10^{-13}$$

The first dissociation is dominant here, and for simple calculations of phosphoric acid in water, we will simply use the first ionization and ignore the other two.

Remember: This is a weak acid. It exists in water mostly as undissociated phosphoric acid molecules.

Solve the equilibrium of phosphoric acid's FIRST proton:



$$K_a = \frac{[\text{H}_2\text{PO}_4^-][\text{H}_3\text{O}^+]}{[\text{H}_3\text{PO}_4]} = 6.9 \times 10^{-3}$$

Species	[Initial]	Δ	[Equilibrium]
H_2PO_4^-	0	+x	x
H_3O^+	0	+x	x
H_3PO_4	0.10	-x	0.10 - x

$$\frac{x^2}{0.10 - x} = 6.9 \times 10^{-3}$$

assume $x \ll 0.10$
so $0.10 - x \approx 0.10$

$$\frac{x^2}{0.10} = 6.9 \times 10^{-3}$$

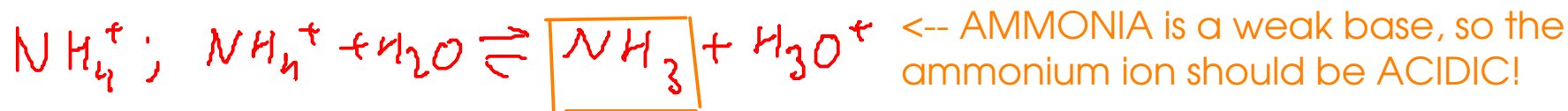
$$x = 0.0262678511 = [\text{H}_3\text{O}^+]$$

$$\text{So, } \text{pH} = -\log(0.0262678511)$$

$$\text{pH} = 1.58$$

163 Find the pH of a solution prepared by dissolving 3.00 g of ammonium nitrate (FW=80.052 g/mol) solid into enough water to make 250. mL of solution.

What's the nature of ammonium nitrate? $\text{NH}_4\text{NO}_3 \rightarrow \text{NH}_4^+ + \text{NO}_3^-$



$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$$

Ka value for ammonium ion? Page A-13 doesn't have it, but page A-14 has Kb for the conjugate, ammonia:

$$K_{b, \text{NH}_3} = 1.8 \times 10^{-5}$$

$$\text{So } K_{a, \text{NH}_4^+} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10}$$

To solve the equilibrium problem, we need to find the nominal concentration of our ammonium nitrate.

$$80.052 \text{ g NH}_4\text{NO}_3 = \text{mol NH}_4\text{NO}_3$$

$$3.00 \text{ g NH}_4\text{NO}_3 \times \frac{\text{mol NH}_4\text{NO}_3}{80.052 \text{ g NH}_4\text{NO}_3} = 0.0374756408 \text{ mol NH}_4\text{NO}_3$$

$$[\text{NH}_4\text{NO}_3]_{\text{nominal}} = \frac{0.0374756408 \text{ mol NH}_4\text{NO}_3}{0.250 \text{ L}} =$$

$$= 0.1499025633 \text{ M NH}_4\text{NO}_3$$



$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} = 5.56 \times 10^{-10}$$

Species	[Initial]	Δ	[Equilibrium]
NH_3	0	+x	x
H_3O^+	0	+x	x
NH_4^+	0.14990	-x	0.14990 - x

Solve ...

$$\frac{x^2}{0.14990 - x} = 5.56 \times 10^{-10}$$

$$\downarrow x \ll 0.14990$$

$$\frac{x^2}{0.14990} = 5.56 \times 10^{-10}$$

$$x = 9.13 \times 10^{-6} = [\text{H}_3\text{O}^+]$$

$$\boxed{\text{pH} = 5.04}$$

... seems reasonable for a weakly acidic salt at moderate concentration.